Chapter-4

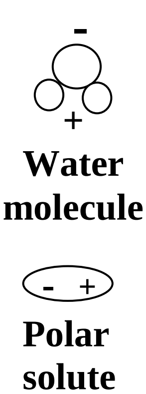
Lesson-1

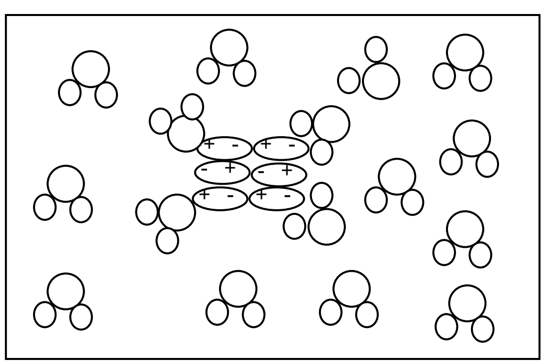
I – General Properties of Aqueous Solutions

- A ***solution*** is a homogeneous mixture.
- A ***solution*** is a mixture of a ***solute*** that is homogeneously distributed through a ***solvent***.
- In a solution, the ***solute particles*** are individual atoms, ions or small molecules surrounded and attracted by the solvent particles.
- In ***aqueous solutions***, the solutes can be classified according to their ability to conduct an electric current through the solution.
- ***Aqueous solutions*** require free moving ions to conduct electricity.
- The solution process is usually ***spontaneous***.

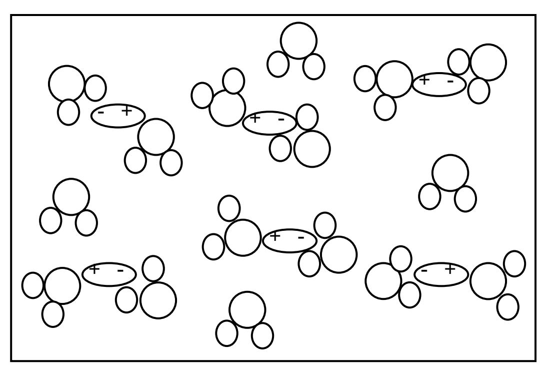
A) Nonelectrolytes:

- A ***nonelectrolyte*** is a solute whose aqueous solution does not conduct electricity because no free moving ions are present.
- ***Nonelectrolytes*** are molecular substances (usually polar) that do not ionize in water.





Alcohols & Sugars



Dipole-dipole forces or hydrogen bonds are formed between the polar molecules and the polar water molecules!

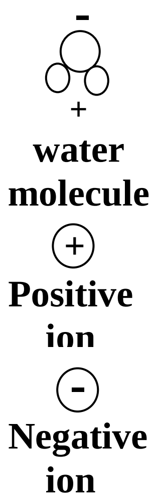
$$C_{12}H_{22}O_{11(s)} \rightarrow C_{12}H_{22}O_{11(aq)}$$

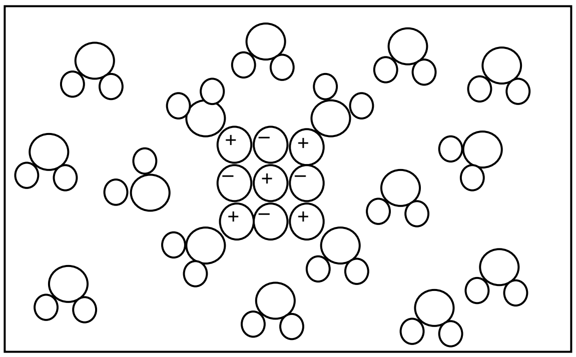
$$C_2H_5OH_{(\ell)} \rightarrow C_2H_5OH_{(aq)}$$

$$I_{2(s)} \rightarrow I_{2(aq)}$$ **(sparingly soluble)**

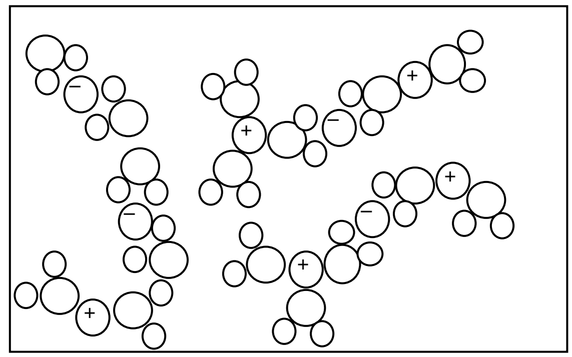
B) Electrolytes:

- An ***electrolyte*** is a solute whose aqueous solution conducts electricity.
- ***Electrolytes*** are ionic substances or also can be molecular substances (usually polar) that ionize in water.

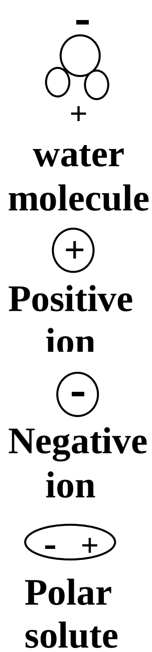


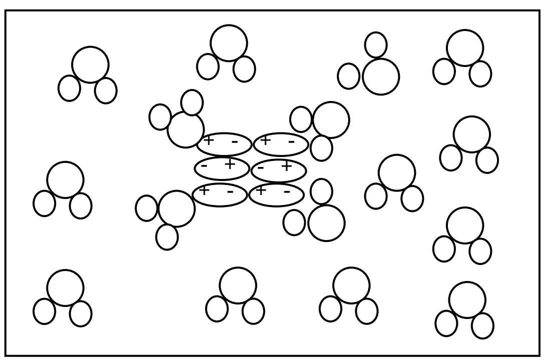


hydroxide bases & salts

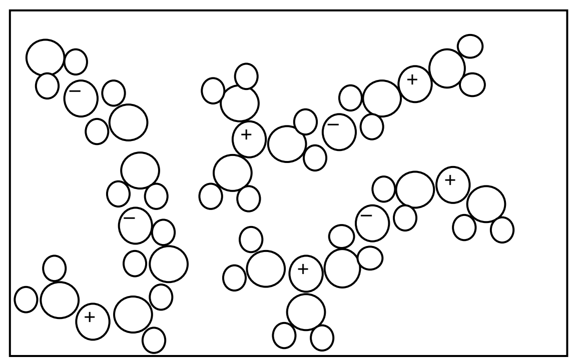


Ion-dipole forces of attraction are formed between dissolved ions and the polar water molecules!





acids & organic bases



Ion-dipole forces of attraction are formed between newly created ions and the polar water molecules!

Ionic electrolytes:

$$Ba(OH)_{2(s)} \rightarrow Ba^{2+}_{(aq)} + 2\,OH^{1-}_{(aq)}$$

$$Na_2CO_{3(s)} \rightarrow 2\,Na^{1+}_{(aq)} + CO_3^{2-}{}_{(aq)}$$

RNH$_2$ is a molecular base.

"R" can be a H atom or a "CH" group.

Molecular electrolytes:

$$H_2SO_{4(s)} + 2\,H_2O_{(\ell)} \rightarrow 2\,H_3O^{1+}_{(aq)} + SO_4^{2-}{}_{(aq)}$$

$$CH_3NH_{2(s)} + H_2O_{(\ell)} \rightarrow CH_3NH_3^{1+}{}_{(aq)} + OH^{1-}_{(aq)}$$

$$NH_{3(g)} + H_2O_{(\ell)} \rightarrow NH_4^{1+}{}_{(aq)} + OH^{1-}_{(aq)}$$

$$CH_3COOH_{(s)} + H_2O_{(\ell)} \rightarrow CH_3COO^{1-}_{(aq)} + H_3O^{1+}_{(aq)}$$

- Smaller or more highly charged ions form stronger **ion-dipole forces of attraction** with water than larger or lower charged ions.

Compound	How to Identify a Compound Based on Its Formula
Acid	• Acid formulas usually start with hydrogen (H). • Organic acid formulas either start with C's and H's and then end with the carboxyl group, (COOH) or start with H and end with $C_xH_yO_z$.
Base	• Base formulas usually start with a metal or a (+) P.A.I. and end with the (-) P.A.I. hydroxide (OH^-). • Organic base formulas can start with elements carbon and hydrogen and then end with the amine group (NH_2). . • Ammonia (NH_3) is a common base.
Salt	• Salt formulas can start with a metal and end with a nonmetal.[†] Salt formulas can start with a (+) P.A.I. and end with a (-) P.A.I.[*] Salt formulas can start with a metal and end with a (-) P.A.I.[*] Salt formulas can start with a (+) P.A.I. and end with a nonmetal.[†]
Sugar	• Sugar formulas will be composed of the three elements: carbon, hydrogen, and oxygen ($C_xH_yO_z$). The elements hydrogen and oxygen • A sugar will **almost always** be in a ratio of 2 to 1(Y:Z = 2:1) but not always. ***Example***: $C_5H_{10}O_5$ is ribose but $C_5H_{10}O_4$ is deoxyribose.
Alcohol	• Alcohol formulas start with elements carbon and hydrogen and then end with the hydroxyl group – OH.

[†](except for oxygen) **[*](except for OH^-)**

Q1: Name several classes of compounds that are classified as electrolytes.

A1: Acids, Bases, and Salts

Q2: Name several classes of compounds that are classified as nonelectrolytes.

A2: Sugars and Alcohols

Q3: Classify the following as an electrolyte or a nonelectrolyte.

(i) CH_3OH (ii) LiI (iii) KOH (iv) $C_6H_{12}O_6$ (v) $HClO_3$

(vi) CH_3NH_2 (vii) NH_4NO_3 (ix) CH_3COOH (x) $C_2H_4(OH)_2$

A3: (i) nonelectrolyte (iv) nonelectrolyte (vii) electrolyte

(ii) electrolyte (v) electrolyte (ix) electrolyte

(iii) electrolyte (vi) electrolyte (x) nonelectrolyte

C) Strong versus Weak Electrolytes:

- ***Strong electrolytes*** are good conductors of electricity in water because they produce lots and lots of free moving ions when dissolved in water.
- ***Strong electrolytes*** include *strong acids, strong bases and soluble salts*.
- **Strong acids**: HI, HBr, HCl, HNO_3, H_2SO_4, $HClO_4$, & $HClO_3$.
- **Strong bases**: LiOH, NaOH, KOH, RbOH, CsOH, and … … $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$.
- **Soluble salts**: Soluble salts are determined by the solubility rules.
- ***Weak electrolytes*** are poor conductors of electricity in water because they produce very small amounts of free moving ions when dissolved in water.
- ***Weak electrolytes*** include *weak acids, weak bases and insoluble salts*.

Q4: Classify the following as a strong electrolyte, a weak electrolyte, or a nonelectrolyte. (i) HF (ii) $Ca(OH)_2$ (iii) HCl (iv) HNO_3 (v) CH_3NH_2 (vi) NaOH (vii) $Mg(OH)_2$ (viii) NH_3 (ix) KOH (x) $HC_2H_3O_2$ (xi) $C_6H_{12}O_6$ (xii) CH_3OH

A4:

(i) weak	(v) weak	(ix) strong
(ii) strong	(vi) strong	(x) weak
(iii) strong	(vii) weak	(xi) non
(iv) strong	(viii) weak	(xii) non

Q5: Write the formulas for *the most common aqueous species* that exist when the following are mixed in water. (i) HF (ii) $Ca(OH)_2$ (iii) HCl (iv) HNO_3 (v) CH_3NH_2 (vi) NaOH (vii) $Mg(OH)_2$ (viii) NH_3 (ix) KOH (x) $HC_2H_3O_2$ (xi) $C_6H_{12}O_6$ (xii) CH_3OH

A5:

(i) HF	(v) CH_3NH_2	(ix) K^{1+} & OH^{1-}
(ii) Ca^{2+} & OH^{1-}	(vi) Na^{1+} & OH^{1-}	(x) $HC_2H_3O_2$
(iii) H^{1+} & Cl^{1-}	(vii) Mg^{2+} & OH^{1-}	(xi) $C_6H_{12}O_6$
(iv) H^{1+} & NO_3^{1-}	(viii) NH_3	(xii) CH_3OH

Chapter-4

Lesson-2

II – Precipitation Reactions

A) Solubilities of Ionic Compounds:

- When ionic solids dissolve in water, the extent to which solution occurs depends on a balance between two forces…
  1) the force of attraction between the water molecules and the ions of the solid (i.e. ***the ion-dipole forces***) and
  2) the force of attraction between the cations and anions of the solid (i.e. ***the ionic bonds***).
- If the ***ion-dipole forces*** dominate, then we would expect the ionic solid to be very soluble in water.
- If the ***ionic bonds*** dominate, then we would expect the ionic solid to be very insoluble in water.
- The solubility of ionic solids covers an enormous range, from $LiClO_3$ (35 mol/L at 25°C) to HgS (10^{-26} mol/L at 25°C)
- By ***learning the basic rules of solubility***, one can predict whether or not an ionic solid will significantly dissolve in water.

<table>
<tr><th>Ions that form Soluble Compounds</th><th>Exceptions</th><th>Ions that form Insoluble Compounds</th><th>Exceptions</th></tr>
<tr><td>Group-1 cations & NH_4^{1+}</td><td></td><td>CO_3^{2-}</td><td>Group-1 cations & NH_4^{1+}</td></tr>
<tr><td>NO_3^{1-}</td><td></td><td>S^{2-}</td><td>Group-1 cations & NH_4^{1+}</td></tr>
<tr><td>$C_2H_3O_2^{1-}$ [CH_3COO^{1-}]</td><td></td><td>PO_4^{3-}</td><td>Group-1 cations & NH_4^{1+}</td></tr>
<tr><td>HCO_3^{1-}</td><td></td><td rowspan="2">CrO_4^{2-}</td><td rowspan="2">Group-1 cations & NH_4^{1+}, & Mg^{2+}, Ca^{2+}</td></tr>
<tr><td>ClO_3^{1-} & ClO_4^{1-}</td><td></td></tr>
<tr><td>Cl^{1-}, Br^{1-}, I^{1-}</td><td>Ag^{1+}, Pb^{2+}, Hg_2^{2+}</td><td rowspan="2">OH^{1-}</td><td rowspan="2">Group-1 cations & NH_4^{1+}, & Ca^{2+}, Sr^{2+}, Ba^{2+}</td></tr>
<tr><td>SO_4^{2-}</td><td>Ag^{1+}, Pb^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}</td></tr>
</table>

B) Molecular, Complete Ionic, and Net-Ionic Equations:

- When solutions of two different ***electrolytes*** (acid, base, or salt) are mixed, an insoluble solid sometimes forms! This solid is called a ***precipitate***.
- The chemical reaction by which an insoluble product is formed is called ***precipitation***.
- A precipitation reaction can be written as a ***molecular equation***.

$$Na_2SO_{4(aq)} + BaCl_{2(aq)} \rightarrow 2\ NaCl_{(aq)} + BaSO_{4(s)}$$

- A precipitation reaction can be written as a ***complete ionic equation***.

$$2\ Na^{1+}_{(aq)} + SO_4^{2-}{}_{(aq)} + Ba^{2+}_{(aq)} + 2\ Cl^{1-}_{(aq)}$$
$$\rightarrow\ 2\ Na^{1+}_{(aq)} + 2\ Cl^{1-}_{(aq)} + BaSO_{4(s)}$$

- A precipitation reaction can be written as a ***net-ionic equation***.

$$SO_4^{2-}{}_{(aq)} + Ba^{2+}_{(aq)} \rightarrow BaSO_{4(s)}$$

- Ions that appear in identical forms among both the products and reactants of a complete ionic equation are called ***spectator ions***.
- If every ion in a complete ionic equation is a spectator ion, then ***no reaction occurs***.

Q1: Write (i) the molecular equation, (ii) the complete ionic equation, and (iii) the net-ionic equation for the reaction that occurs when the two aqueous solutions potassium chromate and aluminum nitrate are mixed together.

A1: (i) $2\,Al(NO_3)_{3(aq)} + 3\,K_2CrO_{4(aq)} \rightarrow 6\,KNO_{3(aq)} + Al_2(CrO_4)_{3(s)}$

(ii) $2\,Al^{3+}_{(aq)} + 6\,NO_3^{1-}{}_{(aq)} + 6\,K^{1+}_{(aq)} + 3\,CrO_4^{2-}{}_{(aq)}$
$\rightarrow\ 6\,K^{1+}_{(aq)} + 6\,NO_3^{1-}{}_{(aq)} + Al_2(CrO_4)_{3(s)}$

(iii) $2\ Al^{3+}_{(aq)} + 3\ CrO_4^{2-}{}_{(aq)} \rightarrow Al_2(CrO_4)_{3(s)}$

Q2: Write (i) the molecular equation, (ii) the complete ionic equation, and (iii) the net-ionic equation for the reaction that occurs when aqueous solutions of silver acetate and sodium iodide are mixed.

A2: (i) $AgC_2H_3O_{2(aq)} + NaI_{(aq)} \rightarrow AgI_{(s)} + NaC_2H_3O_{2(aq)}$

(ii) $Ag^{+}_{(aq)} + C_2H_3O_2{}^{1-}_{(aq)} + Na^{1+}_{(aq)} + I^{1-}_{(aq)}$

$\rightarrow Na^{1+}_{(aq)} + C_2H_3O_2{}^{1-}_{(aq)} + AgI_{(s)}$

(iii) $Ag^{1+}_{(aq)} + I^{1-}_{(aq)} \rightarrow AgI_{(s)}$

Q3: Write (i) the molecular equation, (ii) the complete ionic equation, and (iii) the net-ionic equation for the reaction between the aqueous solutions of lithium sulfate and strontium chlorate.

A3: (i) $Li_2SO_{4(aq)} + Sr(ClO_3)_{2(aq)} \rightarrow 2\ LiClO_{3(aq)} + SrSO_{4(s)}$

(ii) $2\ Li^{1+}_{(aq)} + SO_4{}^{2-}_{(aq)} + Sr^{2+}_{(aq)} + 2\ ClO_3{}^{1-}_{(aq)}$

$\rightarrow 2\ Li^{1+}_{(aq)} + 2\ ClO_3{}^{1-}_{(aq)} + SrSO_{4(s)}$

(iii) $Sr^{2+}_{(aq)} + SO_4{}^{2-}_{(aq)} \rightarrow SrSO_{4(s)}$

Q4: Write (i) the molecular equation, (ii) the complete ionic equation, and (iii) the net-ionic equation for the reaction between the aqueous solutions of ammonium carbonate and calcium nitrate.

A4: (i) $(NH_4)_2CO_{3(aq)} + Ca(NO_3)_{2(aq)} \rightarrow 2\,NH_4NO_{3(aq)} + CaCO_{3(s)}$

(ii) $2\,NH_4{}^{1+}_{(aq)} + CO_3{}^{2-}_{(aq)} + Ca^{2+}_{(aq)} + 2\,NO_3{}^{1-}_{(aq)}$

$\rightarrow 2\,NH_4{}^{1+}_{(aq)} + 2\,NO_3{}^{1-}_{(aq)} + CaCO_{3(s)}$

(iii) $Ca^{2+}_{(aq)} + CO_3{}^{2-}_{(aq)} \rightarrow CaCO_{3(s)}$

<u>Q5</u>: Write (i) the molecular equation, (ii) the complete ionic equation, and (iii) the net-ionic equation for the reaction between the aqueous solutions of sodium hydroxide and iron (III) chloride.

<u>A5</u>: (i) $3\ NaOH_{(aq)} + FeCl_{3(aq)} \rightarrow 3\ NaCl_{(aq)} + Fe(OH)_{3(s)}$

(ii) $3\ Na^{1+}_{(aq)} + 3OH^{1-}_{(aq)} + Fe^{3+}_{(aq)} + 3\ Cl^{1-}_{(aq)}$
$\rightarrow 3\ Na^{1+}_{(aq)} + 3\ Cl^{1-}_{(aq)} + Fe(OH)_{3(s)}$

(iii) $Fe^{3+}_{(aq)} + 3OH^{1-}_{(aq)} \rightarrow Fe(OH)_{3(s)}$

<u>Q6</u>: Write (i) the molecular equation, (ii) the complete ionic equation, and (iii) the net-ionic equation for the reaction that occurs when the aqueous solutions hydrobromic acid and magnesium sulfate are mixed.

<u>A6</u>: (i) $2\ HBr_{(aq)} + MgSO_{4(aq)} \rightarrow H_2SO_{4(aq)} + MgBr_{2(aq)}$

(ii) $2\ H^{1+}_{(aq)} + 2\ Br^{1-}_{(aq)} + Mg^{2+}_{(aq)} + SO_4^{2-}_{(aq)}$
$\rightarrow 2\ H^{1+}_{(aq)} + SO_4^{2-}_{(aq)} + Mg^{2+}_{(aq)} + 2\ Br^{1-}_{(aq)}$

(iii) ***<u>no net ionic reaction occurs</u>***!

IN FACT NO <u>CHEMICAL</u> REACTION OCCURS!

- Reactions of the type:

$$\mathbf{AX + BY \rightarrow AY + BX}$$

  in which cations and anions appear to exchange partners are called ***<u>exchange reactions</u>*** or ***<u>metathesis reactions</u>***.

- ***<u>Precipitation reactions</u>*** and ***<u>acid-base reactions</u>*** usually conform to this pattern.

Chapter-4

Lesson-3

III – Acid-Base Reactions

- In aqueous solutions, HI, HBr, HCl, $HClO_4$, $HClO_3$, HNO_3, and H_2SO_4 are strong acids and exist as free ions.

$\mathbf{H^{1+}_{(aq)} + I^{1-}_{(aq)}}$, $\mathbf{H^{1+}_{(aq)} + Br^{1-}_{(aq)}}$, $\mathbf{H^{1+}_{(aq)} + Cl^{1-}_{(aq)}}$,

$\mathbf{H^{1+}_{(aq)} + ClO_4{}^{1-}_{(aq)}}$, $\mathbf{H^{1+}_{(aq)} + ClO_3{}^{1-}_{(aq)}}$,

$\mathbf{H^{1+}_{(aq)} + NO_3{}^{1-}_{(aq)}}$, $\mathbf{H^{1+}_{(aq)} + HSO_4{}^{1-}_{(aq)}}$

- In aqueous solutions, weak molecular acids exist in equilibrium with their ions, but exist primarily as molecules.

$\mathbf{HF_{(aq)}}$ ℑ $H^{1+}_{(aq)} + F^{1-}_{(aq)}$, $\mathbf{H_3PO_{2(aq)}}$ ℑ $H^{1+}_{(aq)} + H_2PO_4{}^{1-}_{(aq)}$,

$\mathbf{H_2S_{(aq)}}$ ℑ $H^{1+}_{(aq)} + HS^{1-}_{(aq)}$, $\mathbf{HNO_{2(aq)}}$ ℑ $H^{1+}_{(aq)} + NO_2{}^{1-}_{(aq)}$,

$\mathbf{H_2O + CO_{2(aq)}}$ ℑ $H^{1+}_{(aq)} + HCO_3{}^{1-}_{(aq)}$, [$H_2CO_3$]

$\mathbf{H_2O + SO_{2(aq)}}$ ℑ $H^{1+}_{(aq)} + HSO_3{}^{1-}_{(aq)}$ [H_2SO_3]

- In aqueous solutions, LiOH, NaOH, KOH, RbOH, CsOH, $Ba(OH)_2$, $Sr(OH)_2$, and $Ca(OH)_2$, are strong bases and exist as free ions.

$\mathbf{Li^{1+}_{(aq)} + OH^{1-}_{(aq)}}$, $\mathbf{Na^{1+}_{(aq)} + OH^{1-}_{(aq)}}$, $\mathbf{K^{1+}_{(aq)} + OH^{1-}_{(aq)}}$,

$\mathbf{Rb^{1+}_{(aq)} + OH^{1-}_{(aq)}}$, $\mathbf{Cs^{1+}_{(aq)} + OH^{1-}_{(aq)}}$,

$\mathbf{Ba^{2+}_{(aq)} + 2\ OH^{1-}_{(aq)}}$, $\mathbf{Sr^{2+}_{(aq)} + 2\ OH^{1-}_{(aq)}}$, $\mathbf{Ca^{2+}_{(aq)} + 2\ OH^{1-}_{(aq)}}$

- In aqueous solutions, weak molecular bases exist in equilibrium with their ions, but exist primarily as molecules.

$\mathbf{NH_{3(aq)} + H_2O}$ ℑ $NH_4{}^{1+}_{(aq)} + OH^{1-}_{(aq)}$

$\mathbf{CH_3NH_{2(aq)} + H_2O}$ ℑ $CH_3NH_3{}^{1+}_{(aq)} + OH^{1-}_{(aq)}$

$\mathbf{C_2H_5NH_{2(aq)} + H_2O}$ ℑ $C_2H_5NH_3{}^{1+}_{(aq)} + OH^{1-}_{(aq)}$

A) Reactions of Strong Acids with Strong Bases

Q1: Write (i) the molecular equation, (ii) the complete ionic equation, and (iii) the net-ionic equation for the acid-base reaction that occurs when the two aqueous solutions nitric acid and calcium hydroxide are mixed.

A1: (i) $2\ HNO_3 + Ca(OH)_2 \rightarrow Ca(NO_3)_2 + 2\ HOH$

(ii) $2H^{1+} + 2\ NO_3^{1-} + Ca^{2+} + 2\ OH^{1-} \rightarrow Ca^{2+} + 2\ NO_3^{1-} + 2\ HOH$

(iii) $2\ H^{1+} + 2\ OH^{1-} \rightarrow 2\ HOH$

Q2: Write (i) the molecular equation, (ii) the complete ionic equation, and (iii) the net-ionic equation for the acid-base reaction that occurs when the two aqueous solutions hydroiodic acid and potassium hydroxide are mixed.

A2: (i) $HI + KOH \rightarrow KI + HOH$

(ii) $H^{1+} + I^{1-} + K^{1+} + OH^{1-} \rightarrow K^{1+} + I^{1-} + HOH$

(iii) $H^{1+} + OH^{1-} \rightarrow HOH$

Q3: Write (i) the molecular equation, (ii) the complete ionic equation, and (iii) the net-ionic equation for the acid-base reaction that occurs when the two aqueous solutions sulfuric acid and barium hydroxide are mixed.

A3: (i) $H_2SO_4 + Ba(OH)_2 \rightarrow BaSO_4 + 2\ HOH$

(ii) $2\ H^{1+} + SO_4^{2-} + Ba^{2+} + 2\ OH^{1-} \rightarrow BaSO_4 + 2\ HOH$

(iii) $2\ H^{1+} + SO_4^{2-} + Ba^{2+} + 2\ OH^{1-} \rightarrow BaSO_4 + 2\ HOH$

B) Reactions of Weak Acids with Strong Bases

Q4: Write (i) the molecular equation, (ii) the complete ionic equation, and (iii) the net-ionic equation for the acid-base reaction that occurs when **solid** sodium hydroxide is mixed with an aqueous solution of acetic acid.

A4: (i) $HC_2H_3O_2 + NaOH \rightarrow NaC_2H_3O_2 + HOH$

(ii) $HC_2H_3O_2 + NaOH \rightarrow Na^{1+} + C_2H_3O_2^{1-} + HOH$

(iii) $HC_2H_3O_2 + NaOH \rightarrow Na^{1+} + C_2H_3O_2^{1-} + HOH$

Q5: Write (i) the molecular equation, (ii) the complete ionic equation, and (iii) the net-ionic equation for the acid-base reaction that occurs when hydrogen fluoride gas is bubbled through a solution of strontium hydroxide.

A5: (i) $2\ HF + Sr(OH)_2 \rightarrow SrF_2 + 2\ HOH$

(ii) $2\ HF + Sr^{2+} + 2\ OH^{1-} \rightarrow SrF_2 + 2\ HOH$

(iii) $2\ HF + Sr^{2+} + 2\ OH^{1-} \rightarrow SrF_2 + 2\ HOH$

Q6: Write (i) the molecular equation, (ii) the complete ionic equation, and (iii) the net-ionic equation for the acid-base reaction that occurs when a solution of nitrous acid is mixed with a solution of lithium hydroxide.

A6: (i) $HNO_2 + LiOH \rightarrow LiNO_2 + HOH$

(ii) $HNO_2 + Li^{1+} + OH^{1-} \rightarrow Li^{1+} + NO_2^{1-} + HOH$

(iii) $HNO_2 + OH^{1-} \rightarrow NO_2^{1-} + HOH$

C) Reactions of Strong Acids with Weak Bases

Q7: Write (i) the molecular equation, (ii) the complete ionic equation, and (iii) the net-ionic equation for the acid-base reaction that occurs when ammonia gas is bubbled through a solution of hydrochloric acid.

A7: (i) $NH_3 + HCl \rightarrow NH_4Cl$

(ii) $NH_3 + H^{1+} + Cl^{1-} \rightarrow NH_4^{1+} + Cl^{1-}$

(iii) $NH_3 + H^{1+} \rightarrow NH_4^{1+}$

Q8: Write (i) the molecular equation, (ii) the complete ionic equation, and (iii) the net-ionic equation for the acid-base reaction that occurs when a solution of methyl amine is mixed with a solution of perchloric acid.

A8: (i) $CH_3NH_2 + HClO_4 \rightarrow CH_3NH_3ClO_4$

(ii) $CH_3NH_2 + H^{1+} + ClO_4^{1-} \rightarrow CH_3NH_3^{1+} + ClO_4^{1-}$

(iii) $CH_3NH_2 + H^{1+} \rightarrow CH_3NH_3^{1+}$

Q9: Write (i) the molecular equation, (ii) the complete ionic equation, and (iii) the net-ionic equation for the acid-base reaction that occurs when a solution of hydrobromic acid is mixed with a solution of sodium hydrogen carbonate.

A9: (i) $HBr + NaHCO_3 \rightarrow NaBr + H_2O + CO_2$

(ii) $H^{1+} + Br^{1-} + Na^{1+} + HCO_3^{1-} \rightarrow Na^{1+} + Br^{1-} + H_2O + CO_2$

(iii) $H^{1+} + HCO_3^{1-} \rightarrow H_2O + CO_2$

Chapter-4

Lesson-4

IV - Oxidation and Reduction Reactions

A) Changes in Electrons

- During some reactions there is a transfer of electrons from one species to another. These reactions are ***called oxidation-reduction*** reactions or ***redox*** reactions (for short).

$$\mathbf{Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)}}$$

- ***Oxidation*** occurs when a species loses electrons. (**leo**)

$$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2\ e^{1-}$$

The species that undergoes oxidation is said to be ***oxidized***.

"The solid zinc metal is oxidized to zinc ions."

- ***Reduction*** occurs when a species gains electrons. (**ger**)

$$Cu^{2+}_{(aq)} + 2\ e^{1-} \rightarrow Cu_{(s)}$$

The species that undergoes reduction is said to be ***reduced***.

"The copper ions are reduced to solid copper metal."

- During a redox reaction, the oxidation and reduction occur simultaneously!

- During a redox reaction, the number of electrons lost by the species being oxidized must be equal to the number of electrons being gained by the species being reduced.

Q1: For the reaction, $Al_{(s)} + 3\ Ag^{1+}_{(aq)} \rightarrow Al^{3+}_{(aq)} + 3\ Ag_{(s)}$, (i) what species is being oxidized?, (ii) what species is being reduced?, and (iii) how many electrons are being transferred?

A1: (i) $Al_{(s)}$ (ii) Ag^{1+} (iii) 3

B) Oxidation Number

- ***Oxidation number*** is the charge carried by an atom in an ionic species or the "apparent (pseudo) charge" carried by an atom in molecular species.
- Oxidation numbers make it easier to analyze redox reactions.
- Oxidation numbers are assigned to the atoms in a species by the following rules.

  1) The sum of the oxidation numbers for every atom in a species must be equal to the total charge of the species.

  2) The oxidation number for each atom in a free element is zero.

  3) The oxidation number for a monoatomic ion is equal to the charge of the ion.

  4) a) In a compound, the oxidation number for atoms of alkali metals (*group-1 metals*) is +1.

     b) In a compound, the oxidation number for atoms of alkaline earth metals (*group-2 metals*) is +2.

     c) In a compound, the oxidation numbers for atoms of Ag, Zn, Al, and F are +1, +2, +3, and -1 respectively.

  5) In a compound, the oxidation number for atoms of hydrogen is usually +1. In metallic hydrides, the oxidation number for the atoms of hydrogen is -1.

  6) In a compound, the oxidation number for atoms of oxygen is usually -2. In peroxide, the oxidation number for the two oxygen atoms is -1.

  7) In binary compounds in which Group-15, Group-16, and Group-17 elements are the second element, the oxidation numbers are usually -3, -2, and -1 respectively.

  8) In most binary compounds, the atoms of the first element are usually assigned a positive oxidation number and the atoms of the second element are usually assigned a negative oxidation number. [***some noted exceptions***: NH_3, N_2H_4, PH_3, CH_4]

Q2: Give the oxidation numbers for the atoms of each element in the following species:

(a) O_2 (b) $KClO_3$ (c) $Cr_2O_7^{2-}$ (d) H_2O_2 (e) OF_2

(f) KH (g) Hg_2^{2+} (h) $FeSO_4$ (i) NH_3 (j) $AgNO_3$

A2:
(a) O = 0
(b) K = +1, Cl = +5, O = -2
(c) Cr = +6, O = -2
(d) H = +1, O = -1
(e) O = +2, F = -1
(f) K = +1, H = -1
(g) Hg = +1
(h) Fe = +2, S = +6, O = -2
(i) N = -3, H = +1
(j) Ag = +1, N = +5, O = -2

C) Changes in Oxidation Number

- When an atom of a species loses electrons, it will gain oxidation number. Thus oxidation can operationally be defined as a gain in oxidation number.

$$\uparrow\text{oxidation number} = \downarrow\text{electrons}$$

- When an atom of a species gains electrons, it will lose oxidation number. Thus reduction can operationally be defined as a loss in oxidation number.

$$\downarrow\text{oxidation number} = \uparrow\text{electrons}$$

Q3: For the reaction below, which species is oxidized and which species is reduced?

$$\overset{+1\ -1}{HCl}_{(g)} + \overset{+1\ +5\ -2}{HNO}_{3(\ell)} \rightarrow \overset{+4\ -2}{NO}_{2(g)} + \tfrac{1}{2}\overset{0}{Cl}_{2(g)} + \overset{+1\ -2}{H_2O}_{(\ell)}$$

A3: The Cl^{1-} in HCl is oxidized and the N^{+5} in HNO_3 is reduced.

Q4: For the reaction, $2\,KClO_{3(s)} \rightarrow 2\,KCl_{(s)} + 3\,O_{2(g)}$, which species is oxidized and which species is reduced?

A4: The Cl^{+5} in $KClO_3$ is reduced and the O^{-2} in $KClO_3$ is oxidized.

Q5: For the reaction, $P_{4(s)} + 6\,H_2O_{(\ell)} \rightarrow 2\,PH_{3(g)} + 2\,HPO_3{}^{2-}{}_{(aq)} + 4\,H^{1+}{}_{(aq)}$, which species is oxidized and which species is reduced?

A5: The P^0 atoms in P_4 are both oxidized and reduced.

Q6: For the reaction, $2\,Al_{(s)} + 3\,Fe(NO_3)_{2(aq)} \rightarrow 2\,Al(NO_3)_{3(aq)} + 3\,Fe_{(s)}$, which species is oxidized and which species is reduced?

A6: The Al^0 is oxidized and the Fe^{2+} in $Fe(NO_3)_2$ is reduced.

Q7: For the reaction, $Pb_{(s)} + CuSO_{4(aq)} \rightarrow PbSO_{4(s)} + Cu_{(s)}$, which species is oxidized and which species is reduced?

A7: The Pb^0 is oxidized and the Cu^{2+} in $CuSO_4$ is reduced.

Q8: For the reaction, $2\,Li_{(s)} + 2\,HOH_{(\ell)} \rightarrow 2\,LiOH_{(aq)} + H_{2(g)}$, which species is oxidized and which species is reduced?

A8: The Li^0 is oxidized and the H^{1+} in HOH is reduced.

Q9: For the reaction, $Ni_{(s)} + H_2SO_{4(aq)} \rightarrow NiSO_{4(aq)} + H_{2(g)}$, which species is oxidized and which species is reduced?

A9: The Ni^0 is oxidized and the H^{1+} in H_2SO_4 is reduced.

Chapter-4

Lesson-5

V – Displacement Reactions

- A reaction between a metal and an aqueous acid solution or a metal and an aqueous salt solution conforms to the general pattern

$$A + BX \rightarrow AX + B$$

- These reactions are known as ***displacement reactions***.
- Displacement reactions are a type of redox reaction.

A) Reactions of Metals in Aqueous Salt Solutions

- **The pure metal displaces the metal in the salt to produce a new pure metal and a new salt (soluble or insoluble).**

Q1: Write (i) the molecular equation, (ii) the complete ionic equation, and (iii) the net-ionic equation for the redox reaction that occurs when aluminum metal is placed into a solution of iron (II) nitrate.

A1: (i) $2\,Al + 3\,Fe(NO_3)_2 \rightarrow 2\,Al(NO_3)_3 + 3\,Fe$

(ii) $2\,Al + 3\,Fe^{2+} + 6\,NO_3^{1-} \rightarrow 2\,Al^{3+} + 6\,NO_3^{1-} + 3\,Fe$

(iii) $2\,Al + 3\,Fe^{2+} \rightarrow 2\,Al^{3+} + 3\,Fe$

Q2: Write (i) the molecular equation, (ii) the complete ionic equation, and (iii) the net-ionic equation for the redox reaction that occurs when lead metal is placed into a solution of cupric sulfate.

A2: (i) $Pb + CuSO_4 \rightarrow PbSO_4 + Cu$

(ii) $Pb + Cu^{2+} + SO_4^{2-} \rightarrow PbSO_4 + Cu$

(iii) $Pb + Cu^{2+} + SO_4^{2-} \rightarrow PbSO_4 + Cu$

Q3: Write (i) the molecular equation, (ii) the complete ionic equation, and (iii) the net-ionic equation for the redox reaction that occurs when zinc metal is placed into a solution of nickel (II) chloride.

A3: (i) $Zn + NiCl_2 \rightarrow ZnCl_2 + Ni$

(ii) $Zn + Ni^{2+} + 2\,Cl^{1-} \rightarrow Zn^{2+} + 2\,Cl^{1-} + Cu$

(iii) $Zn + Ni^{2+} + \rightarrow Zn^{2+} + Cu$

B) Reactions of Metals in Acids

- **The pure metal displaces the ionizable hydrogen in the acid to produce a salt (soluble or insoluble) and hydrogen gas.**

Q4: Write (i) the molecular equation, (ii) the complete ionic equation, and (iii) the net-ionic equation for the redox reaction that occurs when iron metal is placed hydrochloric acid.

A4: (i) $Fe + 2\,HCl \rightarrow FeCl_2 + H_2$

(ii) $Fe + 2\,H^{1+} + 2\,Cl^{1-} \rightarrow Fe^{2+} + 2\,Cl^{1-} + H_2$

(iii) $Fe + 2\,H^{1+} \rightarrow Fe^{2+} + H_2$

Q5: Write (i) the molecular equation, (ii) the complete ionic equation, and (iii) the net-ionic equation for the redox reaction that occurs when nickel metal is placed sulfuric acid.

A5: (i) $Ni + H_2SO_4 \rightarrow NiSO_4 + H_2$

(ii) $Ni + 2\,H^{1+} + SO_4^{2-} \rightarrow Ni^{2+} + SO_4^{2-} + H_2$

(iii) $Ni + 2\,H^{1+} \rightarrow Ni^{2+} + H_2$

- **Pure metals do not always react spontaneously in all aqueous salt or acid solutions.**
- **The spontaneity of a displacement reaction can be predicted by the use of an activity series of pure metals.**

C) The Activity Series of Pure Metals

- A list of elements arranged in order of decreasing reactivity is called an ***activity series***.
- The activity series for pure metals is below.
- Any pure metal in the series will react spontaneously with (will be oxidized by) the ions of any metal below it in the series.

Most

Activity Series of Pure Metals		
Metals	**Half Reaction**	**Reacts with...**
Li	$Li \rightarrow Li^{1+} + e^-$	* ... cold water, hot water, steam, and acids to produce hydrogen gas
Rb	$Rb \rightarrow Rb^{1+} + e^-$	
K	$K \rightarrow K^{1+} + e^-$	
Cs	$Cs \rightarrow Cs^{1+} + e^-$	
Ba	$Ba \rightarrow Ba^{2+} + 2e^-$	
Sr	$Sr \rightarrow Sr^{2+} + 2e^-$	
Ca	$Ca \rightarrow Ca^{2+} + 2e^-$	
Na	$Na \rightarrow Na^{1+} + e^-$	
Mg	$Mg \rightarrow Mg^{2+} + 2e^-$	** ... steam and acids to produce hydrogen gas
Al	$Al \rightarrow Al^{3+} + 3e^-$	
Ti	$Ti \rightarrow Ti^{2+} + 2e^-$	
Mn	$Mn \rightarrow Mn^{2+} + 2e^-$	
Zn	$Zn \rightarrow Zn^{2+} + 2e^-$	
Cr	$Cr \rightarrow Cr^{3+} + 3e^-$	
Fe	$Fe \rightarrow Fe^{2+} + 2e^-$	
Cd	$Cd \rightarrow Cd^{2+} + 2e^-$	
Co	$Co \rightarrow Co^{2+} + 2e^-$	... acids to produce hydrogen gas
Ni	$Ni \rightarrow Ni^{2+} + 2e^-$	
Sn	$Sn \rightarrow Sn^{2+} + 2e^-$	
Pb	$Pb \rightarrow Pb^{2+} + 2e^-$	
H_2	$H_2 \rightarrow 2H^{1+} + 2e^-$	
Cu	$Cu \rightarrow Cu^{2+} + 2e^-$	Do NOT react with water, steam, or acids to produce hydrogen gas
Ag	$Ag \rightarrow Ag^{1+} + e^-$	
Hg	$Hg \rightarrow Hg^{2+} + 2e^-$	
Pt	$Pt \rightarrow Pt^{2+} + 2e^-$	
Au	$Au \rightarrow Au^{3+} + 3e^-$	

Least

****Little rebellious kids constantly battle strict cantankerous nannies!***

*****Magical aliments titillate manly zones creating feverish cads!***

Q6: Which is more easily oxidized? (i) $Fe_{(s)}$ or $Ni_{(s)}$ (ii) $Pb_{(s)}$ or $Mg_{(s)}$
Which is more easily reduced? (iv) Sn^{2+} or Ti^{2+} (iv) Ag^{1+} or K^{1+}

A6: (i) Fe (ii) Pb (iii) Sn^{2+} (iv) Ag^{1+}

Q7: Determine whether the following pair of reactants will produce a spontaneous reaction. Use **spontaneous** or **nonspontaneous**.

(i) $Au_{(s)} + Cr^{3+}_{(aq)}$ (ii) $Cr_{(s)} + HCl_{(aq)}$ (iii) $Cu_{(s)} + H_2C_2O_{4(aq)}$
(iv) $Ni_{(s)} + CuSO_{4(aq)}$ (v) $H_{2(g)} + Sn^{2+}_{(aq)}$ (vi) $Al_{(s)} + HNO_{3(aq)}$
(vii) $Co_{(s)} + H^{1+}_{(aq)}$ (viii) $Ti_{(s)} + KBr_{(aq)}$ (ix) $H_{2(g)} + AgNO_{3(aq)}$
(x) $Fe_{(s)} + H_2SO_{4(aq)}$ (xi) $Pb_{(s)} + ZnI_{2(aq)}$ (xii) $Ag_{(s)} + Li^{1+}_{(aq)}$

A7:
(i) nonspontaneous
(ii) spontaneous
(iii) nonspontaneous
(iv) spontaneous
(v) nonspontaneous
(vi) spontaneous
(vii) spontaneous
(viii) nonspontaneous
(ix) spontaneous
(x) spontaneous
(xi) nonspontaneous
(xii) nonspontaneous

D) Reactions of Highly Active Metals in Water

- **The pure highly active metal displaces one of the hydrogen atoms in water to produce a hydroxide base and H_2 gas.**

Q8: Write (i) the molecular equation, (ii) the complete ionic equation, and (iii) the net-ionic equation for the redox reaction that occurs when lithium metal is placed into water.

A8:
(i) $2\ Li + 2\ HOH \rightarrow 2\ LiOH + H_2$
(ii) $2\ Li + 2\ HOH \rightarrow 2\ Li^{1+} + 2\ OH^{1-} + H_2$
(iii) $2\ Li + 2\ HOH \rightarrow 2\ Li^{1+} + 2\ OH^{1-} + H_2$

Q9: Write (i) the molecular equation, (ii) the complete ionic equation, and (iii) the net-ionic equation for the redox reaction that occurs when magnesium metal is placed into boiling water.

A9:
(i) $Mg + 2\ H_2O \rightarrow Mg(OH)_2 + H_2$
(ii) $Mg + 2\ H_2O \rightarrow Mg(OH)_2 + H_2$
(iii) $Mg + 2\ H_2O \rightarrow Mg(OH)_2 + H_2$

Chapter-4

Lesson-6

VI – Solution Composition

A) Molarity

- Most reactions take place in solution.
- When an acid solution reacts with a base solution the concentration of at least one of the solutions must be known.
- ***Molarity*** (M) is the most common way that concentration is expressed.

$$\mathbf{M = \frac{moles\ of\ solute}{Liters\ of\ solution}}$$

Q1: What is the molarity of a 125 mL solution with 50.0 g of $C_{12}H_{22}O_{11}$?

A1: $M = \frac{\text{mol of solute}}{\text{liters of soln.}}$ $\qquad 50.0\text{ g } C_{12}H_{22}O_{11} \times \frac{1\text{ mole}}{342\text{ g}} = 0.146\text{ mol}$

$M = \frac{0.146\text{ mol}}{0.125\text{ L}}$

$M = 1.17\text{ M}$

Q2: What is the molarity of a 75.0 mL solution with 15.0 g of KCl?

A2: $15.0\text{ g} \times \frac{1\text{ mol}}{74.6\text{ g}} \times \frac{1}{75.0\text{ mL}} \times \frac{1000\text{ mL}}{1\text{ L}} = 2.68\text{ M}$

Q3: What is the molarity of $\mathbf{Na^{1+}}$ ions & $\mathbf{CO_3^{2-}}$ ions if 10.0 g Na_2CO_3 is dissolved in enough water to make a 100.0 mL solution?

A3: $10.0\text{ g} \times \frac{1\text{ mol}}{106\text{ g}} \times \frac{1}{0.1000\text{ L}} = 0.943\text{ M } Na_2CO_3$

1.89 M Na^{1+} and 0.943 M CO_3^{2-}

Q4: What mass of NaOH is contained in 125 mL of 6.00 M NaOH?

A4: $0.125\ \text{L} \times \frac{6.00\ \text{mol}}{1\ \text{L}} \times \frac{40.00\text{g}}{1\ \text{mol}} = \boxed{30.0\ \text{g NaOH}}$

Q5: What volume (in mL) of 12 M HCl (conc. HCl) must be taken to obtain 0.10 mole of HCl?

A5: $0.10\ \text{mol} \times \frac{1\ \text{L}}{12\ \text{mol}} \times \frac{1000\text{mL}}{1\ \text{L}} = \boxed{8.3\ \text{mL solution}}$

Q6: What volume (in mL) of 0.850 M zinc chloride must be taken to obtain 0.100 mole of chloride ions?

A6: $0.100\ \text{mol}\ Cl^{1-} \times \frac{\mathbf{1\ mol\ ZnCl_2}}{\mathbf{2\ mol\ Cl^{1-}}} \times \frac{1\ \text{L}}{0.850\ \text{mol}} = 0.0588\ \text{L} \rightarrow \boxed{58.8\ \text{mL}}$

Q7: What is the concentration of ions if 10.0 g of NaCl is added to 100.0 mL of 0.125 M Na_2CO_3? *Assume the volume is the same.*

A7: $10.0\ \text{g NaCl} \times \frac{1\ \text{mol}}{58.5\ \text{g}} = 0.171\ \text{mol NaCl}$

$100.0\ \text{mL}\ Na_2CO_3 \times \frac{1\ \text{L}}{1000\ \text{mL}} \times 0.125\ \frac{\text{mol}}{\text{L}} = 0.0125\ \text{mol}\ Na_2CO_3$

***The total moles of Na^{1+} ions = <u>0.196 mol</u>* [*0.171 mol + 2(.0125 mol)*]**

$[Na^{1+}] = \frac{0.196\ \text{mol}}{0.1000\ \text{L}} = \boxed{1.96\ \text{M}}$

$[Cl^{1-}] = \frac{0.171\ \text{mol}}{0.1000\ \text{L}} = \boxed{1.71\ \text{M}}$

$[CO_3^{2-}] = \frac{0.0125\text{mol}}{0.1000\ \text{L}} = \boxed{0.125\ \text{M}}$

Q8: What are the molarities of Na^{1+}, H^{1+}, & Cl^{1-} ions if 10.0 mL of 0.225 M NaCl is added to 15.0 mL of 0.125 M HCl? *Assume volumes are additive.*

A8: $0.0100 \text{ L NaCl} \times 0.225 \frac{\text{mol}}{\text{L}} = 0.00225 \text{ mol NaCl}$

$0.0150 \text{ L HCl} \times 0.125 \frac{\text{mol}}{\text{L}} = 0.00188 \text{ mol HCl}$

$[Na^{1+}] = \frac{0.00225 \text{ mol}}{0.0250 \text{ L}} = \boxed{0.0900 \text{ M}}$

$[H^{1+}] = \frac{0.00188 \text{ mol}}{0.0250 \text{ L}} = \boxed{0.0752 \text{ M}}$

$[Cl^{1-}] = \frac{0.00413 \text{mol}}{0.0250 \text{ L}} = \boxed{0.165 \text{ M}}$

B) Preparing Solutions of Known Molarity

1) Starting with a Solid Solute:

- The steps for preparing a solution using a solid solute are:
  - a) Mass out the solute.
  - b) Put the solute into a volumetric flask and add half the water.
  - c) Swirl the flask to dissolve the solute.
  - d) Add water until the calibration mark is reached.
  - e) Stopper flask and invert to ensure complete mixing.

Q9: Describe how would you prepare a 100.0 mL of a 0.150 M solution of $CuSO_4$ using solid $CuSO_4$?

A9: $100.0 \text{ mL} \times \frac{1\text{L}}{1000 \text{ mL}} \times \frac{0.150 \text{ mol}}{1 \text{ L}} \times \frac{159.6 \text{ g}}{1 \text{ mol}} = 2.39 \text{ g } CuSO_4$

Mass out 2.39 g of $CuSO_4$ and put it into a 100 mL volumetric flask. Add about 50 mL of water and swirl until the $CuSO_4$ was dissolved. Add water to the 100 mL calibration mark. Stopper the flask and invert several times.

2) Starting with a Concentrated Stock Solution:

- In labs, concentrated stock solutions are often prepared for common solutions that are used in various concentrations. It is necessary to dilute the stock solution to the required concentration.
- The dilution equation is:

$$\mathbf{M_{conc} \bullet V_{conc} = M_{dilute} \bullet V_{dilute}}$$

Q10: What would be the molar concentration of a 2.74 M glucose solution if 50.0 ml of it were diluted to 250.0 ml with water?

A10: $M_{conc} \bullet V_{conc} = M_{dilute} \bullet V_{dilute}$

$(2.74)(50.0) = (X)(250.0)$

$X = 0.548$ M

- The steps for preparing a solution using a more concentrated solution are:
  a) Pipette out the needed volume of concentrated solution.
  b) Put this pipetted solution into a volumetric flask.
  c) Add water until the calibration mark is reached.
  d) Stopper flask and invert to ensure complete mixing.

Q11: Describe how would you prepare a 100.0 mL of a 0.15 M solution of HCl from a 12.0 M solution?

A11: $M_{conc} \bullet V_{conc} = M_{dilute} \bullet V_{dilute}$

$(12.0)(X) = (0.15)(100.0)$

$X = 1.25$ mL

Pipette out 1.25 mL of the 12.0 M HCl solution and add it to a 100 mL volumetric flask. Add water to the 100 mL mark. Stopper the flask and invert several times.

Chapter-4

Lesson-7

VII – Solution Stoichiometry - Titrations

- The process of reacting an observed volume of solution of unknown concentration with an observed volume of solution of known concentration (a ***standard solution***) is called ***titration***.
- When conducting a titration, the point at which stoichiometric equivalent quantities of reactants are brought together is known as the ***equivalence point***.
- The ***equivalence point*** is usually signaled by a color change. In titration, the color change is called the ***end point***.

A) Precipitation Titrations

- When pairs of ions, that compose very insoluble compounds, appear in solution together, precipitation occurs.
- If the concentration of one of the ions is known, then the amount of the other ion (necessary for "complete precipitation") can be determined through precipitation titration.

Q1: If 100.0 mL sample of sulfur water ($H_2S_{(aq)}$) was titrated with 16.5 mL of 0.023 M $AgNO_{3(aq)}$ to reach end point, then what was the molar concentration of $H_2S_{(aq)}$ in the water sample?
0.500 g of solid Na_2CrO_4 is added to the 100.0 mL sample to act as an end-point indicator.

Before end-point: $2\ Ag^{1+}_{(aq)} + S^{2-}_{(aq)} \rightarrow Ag_2S_{(s)}$ **black precipitate**

At end-point: $2Ag^{1+}_{(aq)} + CrO_4^{2-}_{(aq)} \rightarrow Ag_2CrO_{4(s)}$ **red precipitate**

A1: $2\ AgNO_{3(aq)} + H_2S_{(aq)} \rightarrow Ag_2S_{(s)} + 2\ HNO_{3(aq)}$

$$0.0165\ L \times 0.023\ \frac{mol}{L}\ AgNO_3 \times \frac{\mathbf{1\ mol\ H_2S}}{\mathbf{2\ mol\ AgNO_3}} \times \frac{1}{0.1000\ L} =$$

$= \boxed{0.0019\ M\ H_2S_{(aq)}}$

Q2: Iron (III) chloride and calcium hydroxide solutions react to form a precipitate. (i) Write the precipitation reaction that occurs. (ii) Determine how many milliliters of 0.200 M $FeCl_3$ is required to completely titrate (react with) 50.0 mL of 0.100 M $Ca(OH)_2$?

0.250 g of solid Na_2CrO_4 is added to the 50.0 mL sample to act as an end-point indicator.

Before end-point: $Fe^{3+}_{(aq)} + 3\,OH^{1-}_{(aq)} \rightarrow Fe(OH)_{3(s)}$ **orange precipitate**

At end-point: $2\,Fe^{3+}_{(aq)} + 3\,CrO_4^{2-}_{(aq)} \rightarrow Fe_2(CrO_4)_{3(s)}$ **red precipitate**

A2: (i) $2\ FeCl_{3(aq)} + 3\ Ca(OH)_{2(aq)} \rightarrow 2\ Fe(OH)_{3(s)} + 3\ CaCl_{3(aq)}$

(ii) $\frac{0.100 \text{ mol } Ca(OH)_2}{L} \times 0.0500\ L \times \frac{\mathbf{2\ mol\ FeCl_3}}{\mathbf{3\ mol\ Ca(OH)_2}} \times \ldots$

$$\ldots \times \frac{1L}{0.200 \text{ mol}} \times \frac{1000 \text{ mL}}{1L} = \boxed{16.7 \text{ mL } FeCl_3}$$

B) Acid-Base Titrations

- Acids react with hydroxide bases to produce a salt and water.

Acid + Base → Salt + Water

- This reaction is called ***neutralization***.
- The point of exact neutralization occurs when the moles of H^{1+} from the acid are equal to the moles of OH^{1-} from the base.

Q3: If it took 25.70 mL of 0.150 M $Ba(OH)_2$ to neutralize 33.40 mL of HCl, then what was the concentration of the 33.40 mL of HCl?

A3: $Ba(OH)_{2(aq)} + 2\ HCl_{(aq)} \rightarrow BaCl_{2(aq)} + 2\ H_2O_{(\ell)}$

$$\frac{0.150 \text{ mol } Ba(OH)_2}{1\ L} \times 0.02570\ L \times \frac{\mathbf{2\ mol\ HCl}}{\mathbf{1\ mol\ Ba(OH)_2}} \times \frac{1}{0.03340\ L} =$$

$$= \boxed{0.231 \text{ M HCl}}$$

Q4: It takes 5.33 mL of 0.150 M nitric acid to neutralize 50.00 mL of a calcium hydroxide solution. (i) Write the neutralization reaction. (ii) Determine the concentration of the calcium hydroxide solution.

A4: (i) $2\ HNO_{3(aq)} + Ca(OH)_{2(aq)} \rightarrow Ca(NO_3)_{2(aq)} + 2\ H_2O_{(\ell)}$

(ii) $$0.150 \frac{\text{mol}}{1\text{ L}} HNO_3 \times 0.00533\text{ L} \times \frac{\mathbf{1\,mol\,Ca(OH)_2}}{\mathbf{2\,mol\,HNO_3}} \times \frac{1}{0.05000\text{ L}} =$$

$$= \boxed{0.00800\text{ M Ca(OH)}_2}$$

Q5: How many milliliters of 0.500 M H_2SO_4 are needed to neutralize 1.00 g of solid NaOH?

A5: $2\ NaOH_{(s)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + 2\ H_2O_{(\ell)}$

$$1.00\text{ g NaOH} \times \frac{1\text{ mol}}{40.0\text{ g}} \times \frac{\mathbf{1\,mol\,H_2SO_4}}{\mathbf{2\,mol\,NaOH}} \times \frac{1\text{ L}}{0.500\text{ mol}} \times \frac{1000\text{ mL}}{1\text{ L}} =$$

$$= \boxed{25.0\text{ mL}}$$

Q6: (i) Write the neutralization reaction of solid oxalic acid [$H_2C_2O_4$] and aqueous potassium hydroxide.
(ii) Determine how many milliliters of 0.100 M potassium hydroxide are needed to fully neutralize 0.875 g of oxalic acid.

A6: (i) $2\ KOH_{(aq)} + H_2C_2O_{4(s)} \rightarrow K_2C_2O_{4(aq)} + 2\ H_2O_{(\ell)}$

(ii) $$0.875\text{ g } H_2C_2O_4 \times \frac{1\text{ mol}}{90.0\text{ g}} \times \frac{\mathbf{2\,mol\,KOH}}{\mathbf{1\,mol\,H_2C_2O_4}} \times \frac{1\text{ L}}{0.100\text{ mol}} = 0.194\text{ L}$$

$$= \boxed{194\text{ mL KOH}}$$

Q7: Write (i) the molecular and (ii) the net ionic neutralization reaction of aqueous acetic acid and aqueous calcium hydroxide.

A7: (i) $Ca(OH)_{2(aq)} + 2\ HC_2H_3O_{2(aq)} \rightarrow Ca(C_2H_3O_2)_{2(aq)} + 2\ H_2O_{(\ell)}$

(ii) $2\ OH^{1-}_{(aq)} + 2\ HC_2H_3O_{2(aq)} \rightarrow 2\ C_2H_3O_2^{1-}{}_{(aq)} + 2\ H_2O_{(\ell)}$

C) Redox Titrations

Q8: If it takes 10.75 mL of 0.0750 M $C_2O_4^{2-}$ to react with 15.00 mL of MnO_4^{1-}, then what is the concentration of the MnO_4^{1-}?

While titrating permanganate with oxalate, the mixture is an intense purple color due to the presence of permanganate ions.

$$5\ \underset{\text{colorless}}{C_2O_4^{2-}{}_{(aq)}} + 2\ \underset{\text{intense purple}}{MnO_4^{1-}{}_{(aq)}} + 16\ H^{1+} \rightarrow 10\ \underset{\text{colorless}}{CO_{2(g)}} + 8\ H_2O_{(\ell)} + 2\ \underset{\text{colorless}}{Mn^{2+}{}_{(aq)}}$$

At the equivalence point, the mixture turns colorless because all the permanganate ions are gone and none of the products produce a color in solution.

A8: $$0.0750\,\frac{\text{mol}}{\text{L}}\ C_2O_4^{2-} \times 0.01075\ \text{L} \times \frac{\mathbf{2\ mol\ MnO_4^{1-}}}{\mathbf{5\ mol\ C_2O_4^{2-}}} \times \frac{1}{0.01500\ \text{L}} =$$

$= \boxed{0.0215\ M\ MnO_4^{1-}}$

Q9: If it takes 7.55 mL of 0.000500 M MnO_4^{1-} to react with 15.00 mL of Fe^{2+}, then what is the concentration of the Fe^{2+}?

$$MnO_4^{1-}{}_{(aq)} + 8\ H^{1+}{}_{(aq)} + 5\ Fe^{2+}{}_{(aq)} \rightarrow 5\ Fe^{3+}{}_{(aq)} + Mn^{2+}{}_{(aq)} + 4\ H_2O_{(\ell)}$$

A9: $$0.000500\,\frac{\text{mol}}{\text{L}}\ MnO_4^{1-} \times 0.00755\ \text{L} \times \frac{\mathbf{5\ mol\ Fe^{2+}}}{\mathbf{1\ mol\ MnO_4^{1-}}} \times \frac{1}{0.01500\ \text{L}} =$$

$= \boxed{0.00126\ M\ Fe^{2+}}$

Chapter-4

Student Activity-2: Titration Problems

Q1: A research chemist isolates a sample of monoprotic nicotinic acid, $HC_6H_4NO_2$ whose molar mass is 123 g/mol. In order to determine its purity, a 0.450 g sample was titrated to end-point with 36.20 mL of 0.100 M NaOH. ***Monoprotic acids have only one ionizable hydrogen.***

(i) Write the molecular neutralization reaction that occurs.
(ii) Determine the number of moles of $HC_6H_4NO_2$ titrated.
(iii) Determine the number of grams of $HC_6H_4NO_2$ in the sample.
(iv) Determine the percent of $HC_6H_4NO_2$ in the sample?
(v) Assuming nicotinic acid follows the "***ate-to-ic***" naming rule, then what is the IUPAC name for $NaC_6H_4NO_2$?

A1: (i) $HC_6H_4NO_{2(aq)} + NaOH_{(aq)} \rightarrow HOH_{(\ell)} + NaC_6H_4NO_{2(aq)}$

(ii) $0.03620 \text{ L} \times 0.100 \frac{\text{mol}}{\text{L}} \text{ NaOH} \times \frac{\mathbf{1\ mol\ HC_6H_4NO_2}}{\mathbf{1\ mol\ NaOH}} =$

$= \boxed{0.00362 \text{ mol } HC_6H_4NO_2}$

(iii) $0.00362 \text{ mol } H_3PO_4 \times \frac{123 \text{ g } HC_6H_4NO_2}{1 \text{ mol}} =$

$= \boxed{0.445 \text{ g } HC_6H_4NO_2}$

(iv) $\% \ HC_6H_4NO_2 = \frac{0.445 \text{ g}}{0.450 \text{ g}} \times 100$

$\% \ HC_6H_4NO_2 = \boxed{98.9 \ \%}$

(v) sodium nicotinate

Q2: The solid nonmetallic oxide - $P_4O_{10(s)}$ in a 4.258 g plant food sample - was converted to $H_3PO_{4(aq)}$ by adding it to 50.00 mL of water. The $H_3PO_{4(aq)}$ was titrated to end-point by 46.90 mL of 0.0820 M $AgNO_{3(aq)}$. ***0.200 g of solid Na_2CrO_4 was added to the 50.00 mL sample to act as an end-point indicator.***

P_4O_{10} conversion: $P_4O_{10(s)} + 4\ H_2O_{(\ell)} \rightarrow 4\ H_3PO_{4(aq)}$

Before end-point: $3\ Ag^{1+}_{(aq)} + PO_4^{3-}{}_{(aq)} \rightarrow Ag_3PO_{4(s)}$ **white precipitate**

At end-point: $2Ag^{1+}_{(aq)} + CrO_4^{2-}{}_{(aq)} \rightarrow Ag_2CrO_{4(s)}$ **red precipitate**

(i) Write the molecular reaction that occurs between $AgNO_3$ and H_3PO_4.
(ii) Determine the number of moles of H_3PO_4 titrated.
(iii) Determine the number of grams of P_4O_{10} in the plant food sample.
(iv) Determine the percent of $P_4O_{10(s)}$ in the plant food sample?

A2: (i) $3\ AgNO_{3(aq)} + H_3PO_{4(aq)} \rightarrow Ag_3PO_{4(s)} + 3\ HNO_{3(aq)}$

(ii) $0.04690\ L \times 0.0820\ \frac{mol}{L}\ AgNO_3 \times \frac{\mathbf{1\ mol\ H_3PO_4}}{\mathbf{3\ mol\ AgNO_3}} =$

$= \boxed{0.00128\ mol\ H_3PO_4}$

(iii) $0.00128\ mol\ H_3PO_4 \times \frac{\mathbf{1\ mol\ P_4O_{10}}}{\mathbf{4\ mol\ H_3PO_4}} \times \frac{284\ g\ P_4O_{10}}{1\ mol\ P_4O_{10}} =$

$= \boxed{0.0909\ g\ P_4O_{10}}$

(iv) $\%\ P_4O_{10(s)} = \frac{0.0909\ g}{4.258\ g} \times 100$

$\%\ P_4O_{10(s)} = \boxed{2.13\ \%}$

Q3: Consider the redox reaction,

$$K_2Cr_2O_{7(aq)} + 7\ H_2SO_{4(aq)} + 6\ FeSO_{4(aq)} \rightarrow$$

$$K_2SO_{4(aq)} + Cr_2(SO_4)_{3(aq)} + 3\ Fe_2(SO_4)_{3(aq)} + 7\ H_2O_{(\ell)}$$

(i) What species is being oxidized?
(ii) What species is being reduced?
(iii) How many moles of electrons are being transferred for each mole of $K_2Cr_2O_7$ reacted?
(iv) If 19.3 mL of $K_2Cr_2O_7$ of unknown concentration is titrated with 25.0 mL of 0.0400 M $FeSO_4$ in order to reach end-point, then what is the molar concentration of the $K_2Cr_2O_7$?
(v) What is the IUPAC name for $K_2Cr_2O_7$?
(vi) What is the IUPAC name for $FeSO_4$?
(vii) What is the traditional name for $FeSO_4$?

A3: (i) Fe^{+2} in the $FeSO_4$

(ii) Cr^{+6} in the $K_2Cr_2O_7$

(iii) 6 moles of electrons

(iv) $$0.0400\ \frac{\text{mol}}{\text{L}}\ FeSO_4 \times 0.0250\ \text{L} \times \frac{\mathbf{1\ mol\ K_2Cr_2O_7}}{\mathbf{6\ mol\ FeSO_4}} \times \frac{1}{0.0193\ \text{L}} =$$

$$= 0.00864\ \text{M}\ K_2Cr_2O$$

(v) potassium dichromate

(vi) iron (II) sulfate

(vii) ferrous sulfate